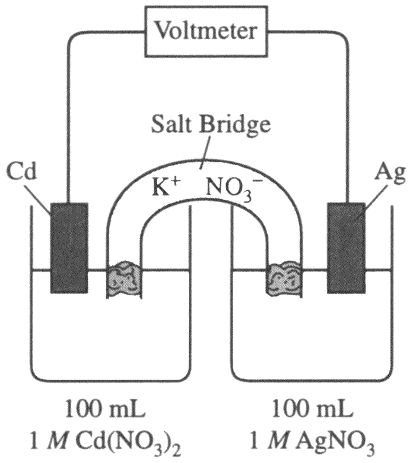
ELECTROCHEMISTRY HMWK



Answer the following questions regarding the electrochemical cell shown.

(a) Write the balanced net-ionic equation for the spontaneous reaction that occurs as the cell operates, and determine the cell voltage.

(b) Determine the value for the equilibrium constant for the reaction in part (a).

(c) Calculate the standard free energy (ΔG^o) for the cell.

(d) What will be the voltage of the cell if 50.0 mL of 2M $Cd(NO_3)_2$ is added to the half-cell on the left? Assume no change in silver cell.

(e) What will be the voltage of the cell if the concentration of the silver ions drops to 0.60 M while the cell runs spontaneously.

(f) What is the function of a salt bridge?

(g) Label the cathode and anode.

(h) Use an arrow to represent the flow of electrons in the wire.

(i) In which direction do anions flow in the salt bridge as the cell operates? Justify your answer.

(j) If 10.0 mL of 3.0-molar $AgNO_3$ solution is added to the half-cell on the right, what will happen to the cell voltage? Explain. No Math

(k) If 1.0 gram of solid NaCl is added to each half-cell, what will happen to the cell voltage? Explain. No Math

(l) If 20.0 mL of distilled water is added to both half-cells, what will happen to the cell voltage. Explain. No Math

In an electrolytic cell, a current of 0.250 ampere is passed through a solution of a chloride of iron, producing $Fe(s)$ and $Cl_2(g)$.

(a) Write the equation for the half-reaction that occurs at the anode.

(b) When the cell operates for 2.00 hours, 0.521 gram of iron is deposited at one electrode. Determine the formula of the chloride of iron in the original solution.

(c) Write the balanced equation for the overall reaction that occurs in the cell.

(d) How many liters of $Cl_2(g)$, measured at 25°C and 750 mm Hg, are produced when the cell operates as described in part (b) ?

(e) Calculate the current that would produce chlorine gas from the solution at a rate of 3.00 grams per hour.